Chapter 10
Outline

__Main Ideas__

I. Kinetic-molecular Theory of Matter
  A. Kinetic-molecular theory of gases
    1. Assumptions of kinetic-molecular theory
      a. Gases consist of large numbers of tiny particles that are far apart relative to their size.
      b. Collisions between gas particles & particles of container walls are perfectly elastic.
        i. Total kinetic energy of particles remains constant at a constant temperature.
      c. Gas particles are in constant, rapid, random motion, therefore possessing kinetic energy.
      d. There are no forces of attraction between gas particles.
      e. The temperature of a gas depends on the average kinetic energy of the gas particles.
        i. $KE = \frac{1}{2} mv^2$ where m = mass & v = velocity
        ii. All gases at same temperature have same kinetic energy, so lighter gases move faster.
    2. Applies only to *ideal gases*
    3. Real gases approach "ideal" at low pressures & high temperatures
  B. Kinetic-molecular theory & the Nature of Gases
    1. Expansion
      a. Gases expand to fill container because particles move in all directions without significant attraction between particles.
    2. Fluidity
      a. Due to insignificant attraction between particles, particles slide past one another like liquids
      b. Because both gases & liquids flow, they're both referred to as *fluids*.
    3. Low Density
      a. Because gas particles are so far apart from each other, density is 1/1000th that of liquid state
    4. Compressibility
      a. Because gas particles are so far apart from each other, volume can be greatly decreased.
    5. Diffusion & Effusion
      a. Because gas particles are in constant random motion, they will mix with other particles
      b. Rates of *effusion* are directly proportional to velocities of their particles
        1. Smaller masses effuse more quickly.
      c. Graham's law of effusion
        1. Rate is inversely proportional to the square root of gas's molar mass.
        2. $Rate_A / Rate_B = \sqrt{molar\ mass_B} / \sqrt{molar\ mass_A}$
    6. Deviation of Real Gases from Ideal Behavior
      a. Particles occupy real space.
      b. Particles exert attractive forces on each other.
      c. Noble gases, especially He & Ne, are most likely to behave "ideally".
      d. Diatomic gases, such as H_2 & N_2, are nonpolar & also "ideal" under certain conditions.
      e. The more polar a gas/ vapor is, such as H_2O & NH_3, the less "ideal" it is.

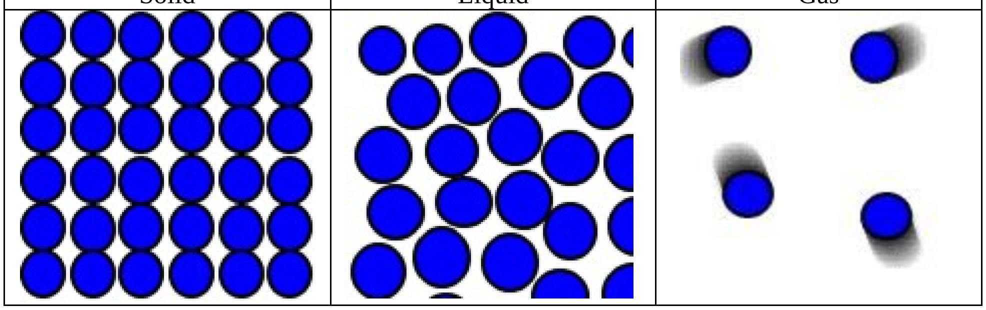
II. Properties of Liquids and the Kinetic Molecular Theory
  A. Fluids
    1. Substances that can flow and therefore take the shape of their container
  B. Relative High Density
    1. 10% less dense than solids (average)
      a. Water is an exception
    2. 1000x more dense than gases
  C. Relative Incompressibility
    1. The volume of liquids doesn't change appreciably when pressure is applied
  D. Ability to Diffuse
    1. Liquids diffuse and mix with other liquids
    2. Rate of diffusion increases with temperature (↑ average Kinetic Energy)
  E. Surface Tension
    1. Surface Tension
      a. A force that tends to pull adjacent parts of a liquid's surface together, thereby decreasing surface area to the smallest possible size
      b. Hydrogen bonding in water creates stronger than normal surface tension
    2. Capillary Action

a. The attraction of the surface of a liquid to the surface of a solid

F. Evaporation and Boiling
1. Vaporization
a. The process by which a liquid of solid changes to a gas
2. Evaporation
a. The process by which particles escape from the surface of a nonboiling liquid enter the gas state
b. Evaporation is a form of vaporization
3. Boiling
a. The change of a liquid to bubbles of vapor that appear throughout the liquid

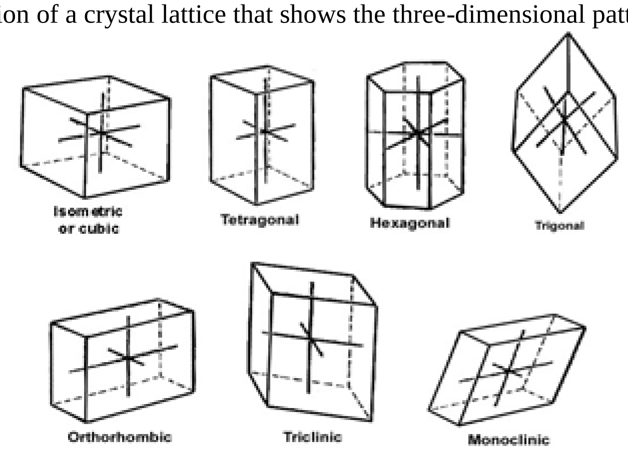
G. Formation of Solids
1. Freezing (or Solidification)
a. The physical change of a liquid to a solid by removal of heat

Solid	Liquid	Gas



III. Properties of Solids and the Kinetic Molecular Theory
A. Types of Solids
1. Crystalline Solids - substances in which the particles are arranged in an orderly, geometric, repeating pattern
2. Amorphous Solids - substances in which the particles are arranged randomly
B. Definite Shape and Volume
C. Definite Melting Point
1. Melting is the physical change of a solid to a liquid by the addition of heat
2. Melting point is the temperature at which a solid becomes a liquid
a. Crystalline solids have definite melting points
b. Amorphous solids do not have definite melting points
D. High Density and Incompressibility
E. Low Rate of Diffusion
1. Two solids in contact will experience VERY SLOW rates of diffusion

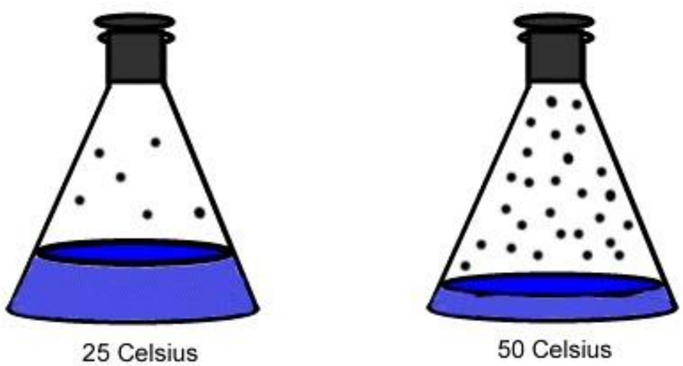
IV. Crystalline Solids
A. Crystal Structure
1. The total three dimensional arrangement of particles of a crystal
B. Unit Cell
1. The smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire lattice



C. Binding Forces in Solids
   1. Ionic crystals
   2. Covalent network crystals
      a. Diamond, quartz
   3. Metallic crystals
   4. Covalent Molecular crystals
      a. Ice

V. Amorphous Solids
   A. "Amorphous"
      1. Greek for "without shape"
   B. Formation of amorphous solids
      1. Rapid cooling of molten materials can prevent the formation of crystals a. Glass
         a. Obsidian

VI. Equilibrium
   A. Equilibrium
      1. Dynamic condition in which two opposing changes occur at equal rates in a closed system
   B. Equilibrium and Changes of State
      1. Phase
         a. Any part of a system that has uniform composition and properties
      2. Condensation
         a. The process by which a gas changes to a liquid
      3. A closed system at constant temperature will reach an equilibrium position at which the rates of evaporation and condensation will be the same
   C. An Equilibrium Equation
      1. liquid + heat enegy $\leftrightarrow$ vapor
   D. Le Chatelier's Principle
      1. When a system at equilibrium is disturbed by application of a stress, it attains a new equilibrium position that minimizes the stress
   E. Equilibrium and Temperature
      1. Increasing the temperature will move the more particles into the vapor phase to compensate for the new energy



   F. Equilibrium and Concentration
      1. If the mass and temperature of a system remain constant, but the volume of the system increases, equilibrium will shift in order to maintain the concentrations of vapor particles

VII. Equilibrium Vapor Pressure of a Liquid
   A. Equilibrium Vapor Pressure
      1. The pressure exerted by a vapor in equilibrium with its corresponding liquid at a given temperature
   B. Volatile Liquids
      1. Liquids that have weak forces of attraction and evaporate easily
   C. Nonvolatile Liquids
      1. Liquids that have strong forces of attraction and do not evaporate easily

VIII. Boiling

Chapter 10
Outline

A. Boiling
   1. The conversion of a liquid to a vapor within the liquid as well as at its surface. It occurs when the equilibrium vapor pressure of the liquid equals the atmospheric pressure

B. Boiling Point
   1. The temperature at which the equilibrium vapor pressure of the liquid equals the atmospheric pressure
      a. Water boils at 100 °C at 1 atm pressure
      b. Water boils above 100 °C at higher pressures
      c. Water boils below 100 °C at lower pressures

C. Molar Heat of Vaporization
   1. The amount of heat energy required to vaporize one mole of a liquid at its boiling point
   2. Strong attractive forces between particles result in high molar heat of vaporization

IX. <u>Freezing and Melting</u>

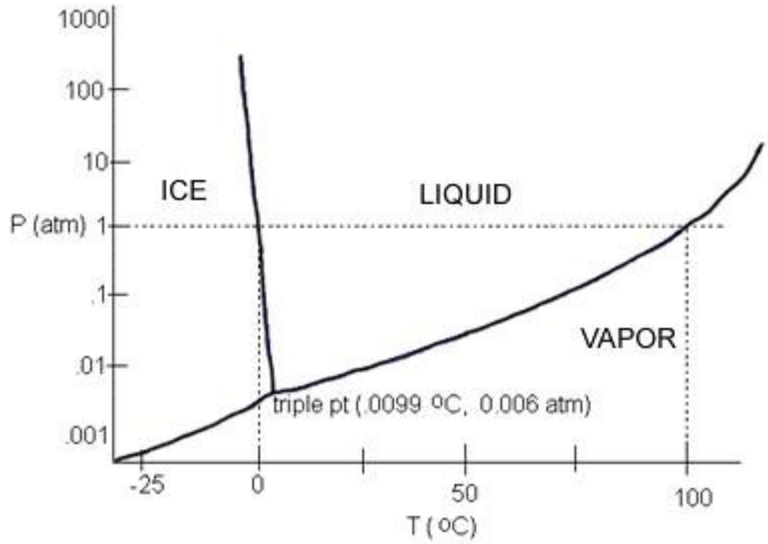
A. Freezing Point
   1. The temperature at which the solid and liquid are in equilibrium at 1 atm
   2. For pure crystalline solids, the melting point and freezing point are the same
   3. Temperature remains constant during a phase change

A. Molar Heat of Fusion
   1. The amount of heat energy required to melt one mole of solid at its melting point

B. Sublimation and Deposition
   1. Sublimation is the change of state from a solid directly to a gas
      a. Dry ice → Gaseous CO_2
   2. Deposition is the change of state from a gas directly to a solid

X. <u>Phase Diagrams</u>

A. Phase Diagram
   1. A graph of pressure versus temperature that shows the conditions under which the phases of a substance exist (notice that pressure is on a logarithmic scale)



B. Triple Point
   1. The temperature and pressure conditions at which the solid, liquid, and vapor of the substance can coexist at equilibrium

C. Critical Temperature
   1. The temperature at above which the substance cannot exist in the liquid state, regardless of pressure
      a. For water, the critical temperature is 373.99 °C

D. Critical Pressure
   1. The lowest pressure at which the substance can exist as a liquid at the critical temperature
      a. For water, the critical pressure is 217.75 atm

E. Critical Point
   1. The point on the graph describing simultaneously the critical temperature and the critical pressure

P = 217.75 atm Temperature = 373.99 °C

Bond Type	Polar		Density of ice (0 °C)	0.917 g/cm^3
Bond angle	105°		Density of water (0 °C)	0.999 g/cm^3
Boiling point	100 °		Point of maximum density	3.98 °C
Melting Point	0 °C		Molar heat of fusion	6.009 kJ/mole
			Molar heat of vaporization	40.79 kJ/mole